Electromagnetic Radiation

- **Carries energy through space (also known as radiant energy).**
- **Includes visible light, x-rays, radio waves, heat radiation from a fireplace**
- **Share certain fundamental characteristics**
- **All move through a vacuum at a speed of 3.00 x 10^8 m/s, the "speed of light."**
- **Have "wave-like" characteristics .**

The Wave Nature of Light

Most of our present understanding of the electronic structure of atoms has come from analysis of the light emitted or absorbed by substances.

- **https://www.youtube.com/watch?v=1flI4UfNFKk**

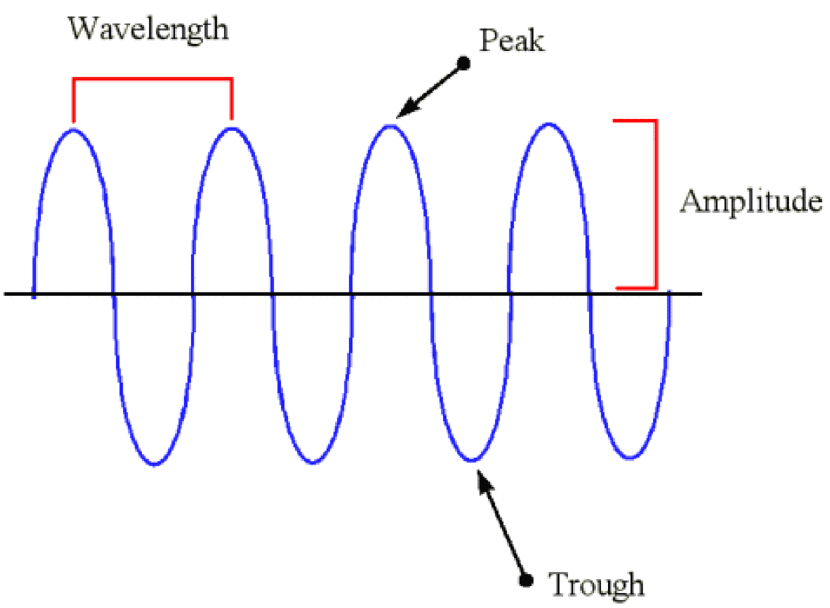
Electrons hold the key to understanding why substances behave as they do.

***When atoms react it is their outer pars, their electrons, that interact*.**

We refer to the arrangements of electrons in atoms as *their electronic structure*.

- **Number of electrons**
- **Where they can be found**
- **The energies they possess**

The number of complete wavelengths, or *cycles*, that pass a given point in 1 second is the frequency of the wave . (frequency=cycles/second)

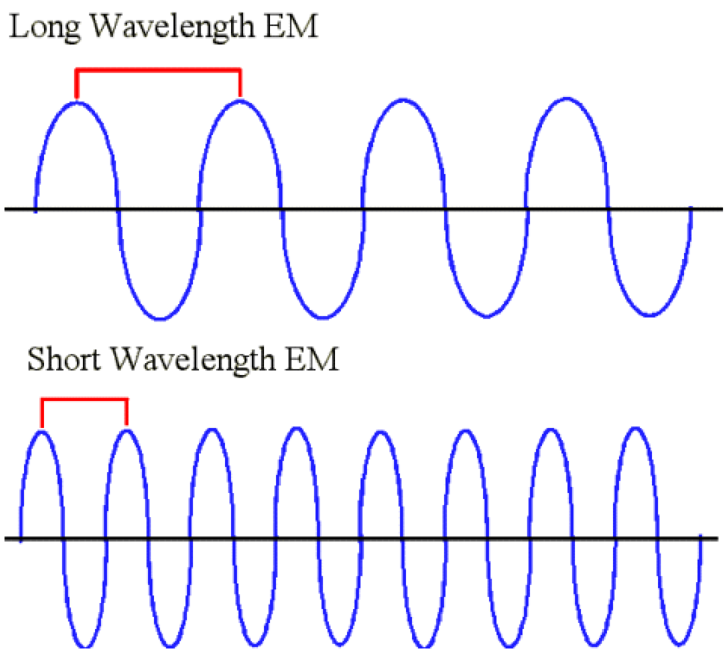


- **https://www.youtube.com/watch?v=m4t7gTmBK3g**

Electromagnetic radiation has both *electric* and *magnetic* properties. The wave-like property of electromagnetic radiation is due to the periodic oscillations of these components.

We can assign a *frequency* and a *wavelength* to electromagnetic radiation

Because all electromagnetic radiation moves at the same speed (speed of light) *wavelength and frequency are related*



- **If the *wavelength is long*, there will be fewer cycles passing a given point per second, thus *the frequency will be low***
- **If the *wavelength is short*, there will be more cycles passing a given point per second, and the *frequency will be high***

Thus, there is *an inverse relationship between wavelength and frequency*

$$\text{frequency} = \left(\frac{1}{\text{wavelength}}\right) * \text{speed of light}$$

$$\nu = \left(\frac{1}{\lambda}\right) * c$$

$$\nu\lambda = c$$

- **(*frequency [ν] * wavelength [λ]* is a constant (*c*)**

What is the speed of a wave?

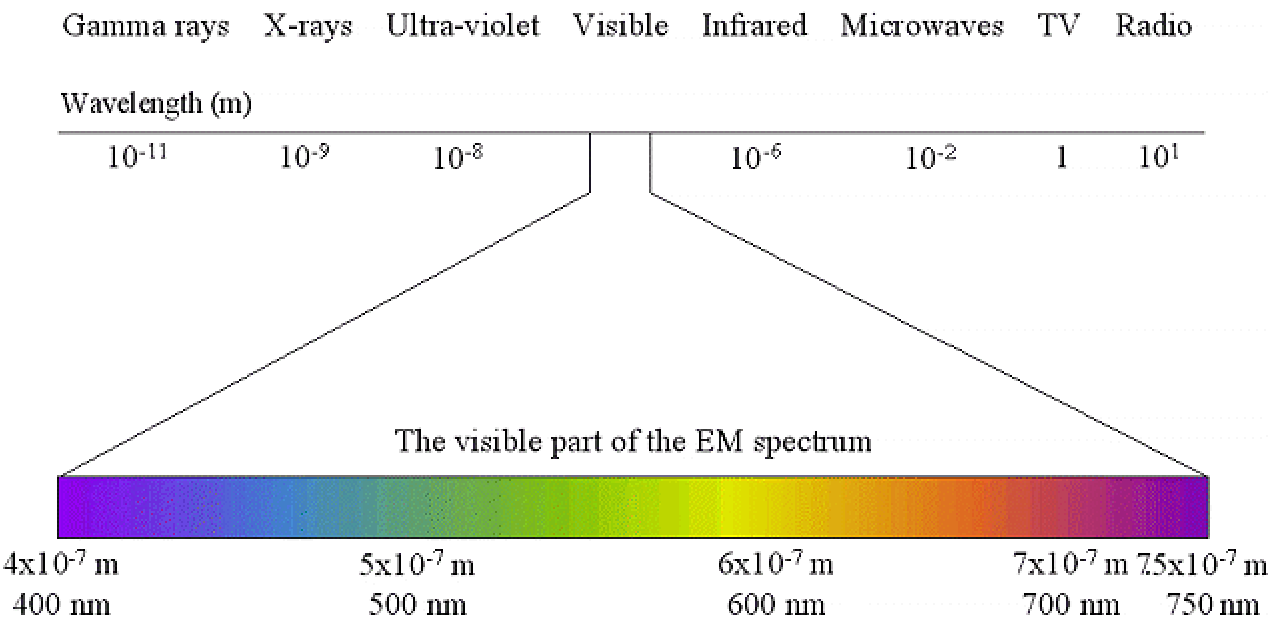
Imagine you are on the beach watching the ocean waves go by, and you want to know the speed of the waves. There is an island offshore with a palm tree that will serve as a convenient frame of reference. You count the number of waves that pass by the tree in one minute:

In this case, two peaks (two wavelengths) pass by the tree in one minute. Thus, the frequency is 2 wavelengths/minute.

If we measure the distance between the peaks (i.e. the wavelength) we can determine the speed of the wave:

Speed of the wave = (distance between peaks) * (frequency)

= (wavelength) * (frequency)



The unit of length chosen to describe a particular wavelength is typically dependent on the type of electromagnetic radiation

Unit	Symbol	Length (m)	Type of Radiation
Angstrom	**Å**	**10^{-10}**	**X-ray**
Nanometer	**nm**	**10^{-9}**	**UV, visible**
Micrometer	**m**	**10^{-6}**	**Infrared**
Millimeter	**mm**	**10^{-3}**	**Infrared**
Centimeter	**cm**	**10^{-2}**	**Microwave**
Meter	**m**	**1**	**TV, radio**

The range of EM wavelengths is dramatic

- The wavelengths of *gamma-rays* (<0.1 Å) are similar to the diameter of *atomic nuclei*
- The wavelengths of some *radio waves* can be larger than *a football field*
- https://www.youtube.com/watch?v=EEgAWW4vuiI
- https://www.youtube.com/results?search_query=the+wave+nature+of+light+chemistry

Frequency

- Frequency is expressed in *cycles per second*, also known as hertz (Hz)
- Usually the dimension 'cycles' is omitted and *frequencies thus have the dimension of* s^{-1}

Sodium vapor lamps are sometimes used for public lighting. They give off a yellowish light with a wavelength of 589 nm. What is the frequency of this radiation?

frequency*wavelength = speed of light

frequency = speed of light/wavelength

$$= (3.00 \times 10^{8}\ m/s)/(589 \times 10^{-9} m)$$

$$= 5.09 \times 10^{14}\ s^{-1}$$

$= 5.09 \times 10^{14}$ cycles per second or 5.09×10^{14} hertz

Electronic structure of atoms quantum effects and photons

What's the difference between a "red hot" poker and a "white hot" poker?

- Poker is a metal rod with a handle, used for prodding and stirring an open fire.
- The pokers are different temperatures ("white hot" poker has a higher temperature)
- The pokers emit different intensities and wavelengths of electromagnetic radiation (especially in the visible spectrum).
- https://www.youtube.com/watch?v=eu3qvI8oxdU

Max Planck (1900)

Energy can be released (or absorbed) by atoms only in "packets" of some minimum size

- This minimum energy packet is called a *quantum*
- The energy (E) of a quantum is related to its frequency (n) by some constant (h): $E = h\nu$

- ***h*** **is known as "*Planck's constant*", and has a value of 6.63×10^{-34} Joule seconds (Js)**
- ***Electromagnetic energy is always emitted or absorbed in whole number multiples of (h* ν)***

■ Example : Calculate the smallest amount of energy (i.e. one quantum) that an object can absorb from yellow light with a wavelength of 589 nm

$$\text{Energy quantum} = h\,\nu$$

so we need to know the frequency ν

$$\nu\,\lambda = c$$

$$\nu = c / \lambda$$

$$\nu = (3.00 \times 10^{8}\ \text{m/s}) / (589 \times 10^{-9}\ \text{m})$$

$$\nu = 5.09 \times 10^{14}\ \text{s}^{-1}$$

plugging into Planck's equation: $E = h * \nu$

$$E = (6.63 \times 10^{-34}\ \text{Js})*(5.09 \times 10^{14}\ \text{s}^{-1})$$

$$E\ (1\ \text{quanta}) = 3.37 \times 10^{-19}\ \text{J}$$

Note that a quanta is quite small.

When we receive infrared radiation from a fireplace we absorb it in quanta according to Planck's Law.

However, we can't detect that the energy absorption is incremental. تدريجي

The Photoelectric Effect

- **Light shining on a metallic surface can cause the surface to *emit* electrons**
- **For each metal there is *a minimum frequency of light below which no electrons are emitted*, regardless of the *intensity* of the light**
- **The higher the light's frequency above this minimum value, the greater the kinetic energy of the released electron(s)**
- https://www.youtube.com/watch?v=muxRZ1irsrk

Using Planck's results Einstein (1905) was able to deduce the basis of the photoelectric effect

- **Einstein assumed that the light was a stream of tiny energy packets called *Photons***

- **Each photon has an energy proportional to its frequency ($E=h\nu$)**
- **When a photon strikes the metal its energy is transferred to an electron**
- **A certain amount of energy is needed to overcome the attractive force between the electron and the protons in the atom**

Thus, if the quanta of light energy absorbed by the electron is insufficient for the electron to overcome the attractive forces in the atom, the electron will not be ejected - regardless of the intensity of the light.

If the quanta of light energy absorbed is greater than the energy needed for the electron to overcome the attractive forces of the atom, then the excess energy becomes kinetic energy of the released electron.

Since different metals have different atomic structure (number of protons, different electronic structure) the quanta of light needed to overcome the attractive forces within the atom differs for each element.

The photoelectric effect :

The electron will be released from the atom if it absorbs a photon with enough energy.

Einstein's interpretation of the photoelectric effect suggests that light has characteristics of particles. Is light a wave or does it consist of particles?

Electronic Structure of Atoms

Bohr's model of the hydrogen atom

In 1913 Niels Bohr developed a theoretical explanation for a phenomenon known as *line spectra*.

Bohr's Model of the Hydrogen Atom

- **https://www.youtube.com/watch?v=in1j2L7Kf_8**

Line Spectra

Lasers emit radiation which is composed of a *single wavelength*. However, most common sources of emitted radiation (i.e. the sun, a lightbulb) produce radiation containing *many different wavelengths*.

When the different wavelengths of radiation are separated from such a source a *spectrum* is produced.

- **https://www.youtube.com/watch?v=YAAw96Kex-I**

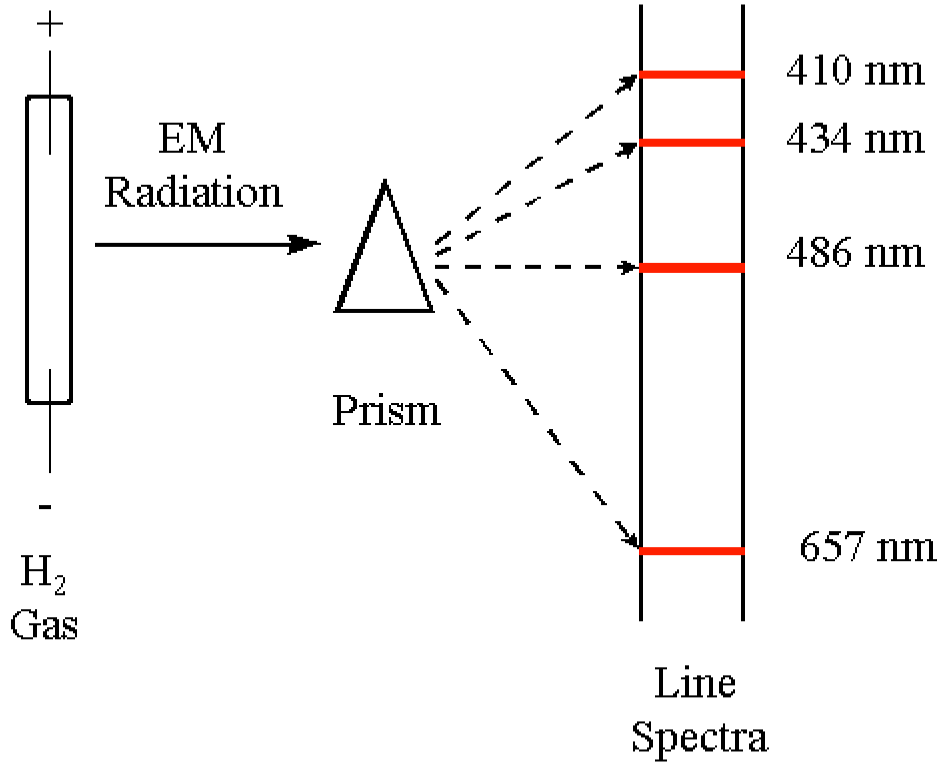
A *rainbow*

- **represents the spectrum of wavelengths of light contained in the light emitted by the sun**
- **Sun light passing through a *prism* (or raindrops) is separated into its component wavelengths**
- **Sunlight is made up of a *continuous spectrum* of wavelengths (from red to violet) - there are no gaps**

Not all radiation sources emit a continuous spectrum of wavelengths of light

- **When high voltage is applied to a glass tube containing various gasses under low pressure different colored light is emitted**
  - **Neon gas produces a red-orange glow**
  - **Sodium gas produces a yellow glow**
- **When such light is passed through a prism only a *few wavelengths* are present in the resulting spectra**
  - **These appear as lines separated by dark areas, and thus are called *line spectra***

When the spectrum emitted by *hydrogen gas* was passed through a prism and separated into its constituent wavelengths *four lines* appeared at characteristic wavelengths .



Bohr's Model

- **Bohr began with the assumption that electrons were orbiting the nucleus, much like the earth orbits the sun.**
- **From classical physics, a charge traveling in a circular path should lose energy by emitting electromagnetic radiation**
- **If the "orbiting" electron loses energy, it should end up spiraling into the nucleus (which it does not). *Therefore, classical physical laws either don't apply or are inadequate to explain the inner workings of the atom***
- **Bohr borrowed the idea of quantized energy from Planck**
  - **He proposed that only orbits of certain radii, corresponding to defined energies, are "permitted"**
  - **An electron orbiting in one of these "allowed" orbits:**
    - **Has a defined energy state**
    - **Will not radiate energy**
    - **Will not spiral into the nucleus**

If the orbits of the electron are restricted, the energies that the electron can possess are likewise restricted and are defined by the equation:

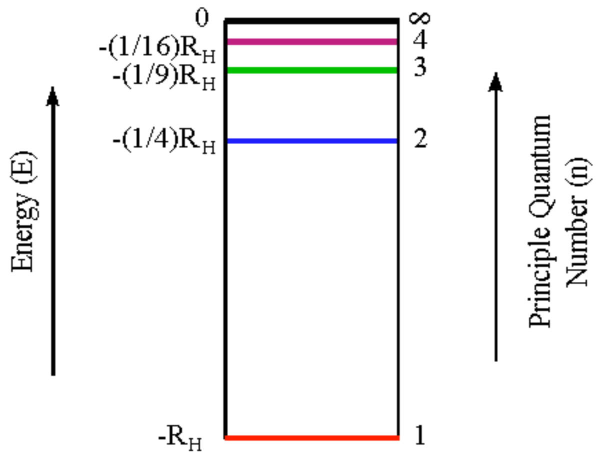
$$E_n = (-R_H)\left(\frac{1}{n^2}\right)$$

Where R_H is a constant called the *Rydberg constant* and has the value 2.18 x 10^{-18} J

***'n'* is an integer, called the *principle quantum number* and corresponds to the different allowed orbits for the electron.**

Thus, an electron in the first allowed orbit (closest to the nucleus) has *n=1*, an electron in the next allowed orbit further from the nuclei has *n=2*, and so on.

Thus, the relative energies of these allowed orbits for the electrons can be diagrammed as follows:



All the relative energies are *negative*

- **The lower the energy, the more stable the atom**
- **The lowest energy state (n=1) is called the *ground state* of the atom**
- **When an electron is in a higher (less negative) energy orbit (i.e. n=2 or higher) the atom is said to be in an *excited state***
- **As n becomes larger, we reach a point at which the electron is completely separated from the nucleus**
  - **E = (-2.18 x 10^{-18} J)(1/infinity) = 0**
  - ***Thus, the state in which the electron is separated from the nucleus is the reference or zero energy state* (actually *higher* in energy than other states)**

Bohr also assumed that the electron can change from one allowed orbit to another

- **Energy must be absorbed for an electron to move to a higher state (one with a higher n value)**
- **Energy is emitted when the electron moves to an orbit of lower energy (one with a lower n value)**
- **The overall change in energy associated with "orbit jumping" is the difference in energy levels between the ending (final) and initial orbits:**

$$\Delta E = E_f - E_i$$

$$E_n = (-R_H)\left(\frac{1}{n^2}\right)$$

Substituting in for the previously defined energy equation:

$$\Delta E = \left(\frac{-R_H}{n_f^2}\right) - \left(\frac{-R_H}{n_i^2}\right) = (-R_H) * \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = R_H * \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

When an electron "falls" from a higher orbit to a lower one the energy difference is a defined amount and results in emitted electromagnetic radiation of a defined energy (ΔE)

- **Planck had deduced that the energy of the photons comprising EM radiation is a function of its frequency (E = h ν)**
- **Therefore, if the emitted radiation from a falling electron had a defined energy, then it must have a correspondingly defined frequency**

$$\Delta E = R_H * \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) = h\nu$$

<u>Note:</u>

- ***ΔE is positive*** **when n_f is greater than n_i, this occurs when energy is absorbed and an electron moves up to a higher energy level (i.e. orbit).**
- ***When ΔE is negative, radiant energy is emitted and an electron has fallen down to a lower energy state***

Revisiting Balmer's equation:

In 1885 a Swiss school teacher figured out that the *frequencies* of the light corresponding to these <u>wavelengths</u> fit a relatively simple mathematical formula:

$$\nu = C*\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \text{ where } n = 3, 4, 5, 6$$

where C = 3.29 x 10^{15} s^{-1} (*not the 'c' used for the speed of light*)

Since energy lost by the electrons is energy "gained" by the emitted EM energy, the EM energy from Bohr's equation would be:

$$h\nu = R_H*\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

Thus, Balmer's constant 'C' = (R_H/h) (Rydberg constant divided by Planck's constant), and n_f = 2.

Thus, the only emitted energies which fall in the visible spectrum are from those electrons which fell down to the *second* quantum orbital.

Those which fell down to the first orbital have a higher energy (frequency) than can be seen in the visible spectrum.

■ Example : Calculate the wavelength of light that corresponds to the transition of the electron from the n=4 to the n=2 state of the hydrogen atom. Is the light absorbed or emitted by the atom?

Since the electron is "falling" from level 4 down to level 2, energy will be given up and manifested as emitted electromagnetic radiation:

$$\Delta E = R_H*\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) = h\nu$$

ΔE = (2.18 x 10^{-18} J)((1/16)-(1/4)) = -4.09 x 10^{-19} J (light is emitted)

4.09 x 10-19 J = (6.63 x 10^{-34} Js) * (ν)

$$6.17 \times 10^{14}\ s^{-1} = \nu$$

$$\lambda = (3.00 \times 10^{8}\ m\ s^{-1})/\ (6.17 \times 10^{14}\ s^{-1}) = 4.87 \times 10^{-7} m = 487\ nm$$

Bohr's model of the atom was important because it introduced quantized energy states for the electrons.

However, as a model it was only useful for predicting the behavior of atoms with a single electron (H, He^{+}, and Li^{2+} ions).

Thus, a different model of the atom eventually replaced Bohr's model. However, we will retain the concept of quantized energy states

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